HISTORY OF ATOMIC THEORY NOTES

I. Ancient Greeks
   A. Aristotle
      ~ believed there were 4 elements:
   B. Democritus
      ~ believed in the "particle theory":
      ~ named the small particles "atoms" which means

II. Dalton (1808)
   A. Father of Atomic Theory
   B. Matter is composed of atoms
   C. Points to his theory
      ~ atoms of an element
      ~ law of conservation of matter
      ~ Law of Multiple Proportions

III. Thomson (1897)
   A. discovered the
   B. used a

IV. Millikan (1909)
   A. discovered electrons have
   B. his experiment called

V. Rutherford (1911)
   A. discovered
   B. his experiment called
   C. other important aspects of his experiment

VI. Bohr (1913)
   A. electrons
   B. his model called
   C. energy must be

VII. Chadwick (1932)
   A. discovered

VIII. Progression of the atomic model

Democritus/Dalton --> Thomson --> Rutherford --> Bohr --> Quantum Mechanical

SUBATOMIC PARTICLES NOTES

Three types of subatomic particles:

Particle	Symbol	Charge	Location	Mass (kg)	Mass (amu)
proton					
neutron					
electron					

Parts of the atom:

- NUCLEUS: accounts for almost all of the _____________________ of an atom
- ELECTRON CLOUD: accounts for almost all of the _____________________ of an atom

WEIGHING ATOMS NOTES

ISOTOPES:

NUCLIDE:

To differentiate between the different nuclides, there are several different notations that are used:

$^{A}_{Z}E$ ^{A}E Element - A

A = mass number
Z = atomic number
E = symbol for element
"Element" = name of element
Mass number = number of protons plus the number of neutrons in the nucleus of that nuclide
= will always be a WHOLE number
= rounded decimal number on Periodic Table is the mass number of the MOST COMMON isotope - not ALL isotopes of that element!
Atomic number = number of protons in the nucleus
= whole number on the Periodic Table
So, if mass number = # of protons + # of neutrons, and atomic number = # of protons...
then # of neutrons =

ATOMIC CHART 1 WORKSHEET

	Symbol	Name	Atomic #	Mass #	# of protons	# of neutrons	# of electrons
1.)		calcium - 40					
2.)	$^{201}_{80}Hg$						
3.)			53	127			
4.)					19	20	
5.)				40		22	
6.)	$^{207}_{82}Pb^{+2}$						
7.)		barium - 137					
8.)			14	28			
9.)				19		10	
10.)			26			30	

ATOMIC CHART 2 WORKSHEET

	Symbol	Name	# of p^+	# of n^o	Atomic #	Mass #	# of e^-	Period	Group
1.)	$^{14}N^{-3}$								
2.)	$^{27}Al^{+3}$								
3.)			47			108			
4.)				45		80			
5.)					55	134			
6.)				18	17				
7.)			92			238			
8.)					38	86			
9.)		phosphorus - 31							
10.)	$^{87}Rb^{+1}$								
11.)						75		4	15
12.)				118				6	11

ISOTOPES AND AVERAGE ATOMIC MASS NOTES

ISOTOPES: atoms of the same element that have different masses

Q: What particles make up the mass of an atom?

A: protons (p^+) and neutrons (n^o)

So, isotopes of the same element have the same number of _______________, but have different numbers of _______________.

MASS NUMBER: # of p^+ plus # of n^o

Q: Why is the mass given on the Periodic Table a decimal number (not a whole #)?

A: isotopes of elements occur in nature a certain % of the time. Decimal # on the P.T. (called the "AVERAGE ATOMIC MASS") is the weighted average of all isotopes of that element

EXAMPLE:

Chlorine exists in nature as two different isotopes. Chlorine - 35 has a natural abundance of 75.8 %. Chlorine - 37 has a natural abundance of 24.2 %. What is the average atomic mass of chlorine? (On these problems, you MUST show your work in order to receive any credit.)

^{35}Cl: 35 amu x 0.758 = 26.53

^{37}Cl: 37 amu x 0.242 = 8.954

26.53 + 8.954 = 35.484 amu

TRY THIS ONE ON YOUR OWN FOLLOWING THE EXAMPLE ABOVE:

An imaginary element with the symbol Lb has 3 naturally-occurring isotopes. ^{41}Lb has a percent abundance of 41.0 %. ^{42}Lb has a percent abundance of 48.3 %. ^{43}Lb has a percent abundance of 10.7 %. What is the average atomic mass of element Lb?

AVERAGE ATOMIC MASS WORKSHEET

Determine the average atomic mass of the following mixtures of isotopes.

1.) 75 % ^{133}Cs, 20 % ^{132}Cs, and 5 % ^{134}Cs

2.) 80 % ^{127}I, 17 % ^{126}I, 3 % ^{128}I

3.) 50 % ^{197}Au, 50 % ^{198}Au

4.) 15 % ^{55}Fe, 85 % ^{56}Fe

5.) 99 % ^{1}H, 0.8 % ^{2}H, 0.2 % ^{3}H

6.) 95 % ^{14}N, 3 % ^{15}N, 2 % ^{16}N

7.) 98 % ^{12}C, 2 % ^{14}C

COUNTING ATOMS NOTES

When we want to know how many atoms of a substance are in a sample of the substance that we can see, counting the atoms individually is not practical. First of all, atoms are so small that we cannot see them individually with our eyes. Second, we would be counting for a very, very, very long time. Here is an example for you to think about. The "lead" of a pencil is made of pure carbon (graphite). Think about the size of the dot on a "i" that you would write on your paper with a pencil. Not very big. The number of atoms of carbon in that dot is roughly equal to the population of the world - about 6 billion! It would take a very long time to count that high. And that is just for the number of atoms in the dot on an "i" on your paper!

So, you can see that we have a problem if we try to count atoms individually. Instead, chemists have come up with a system that is much easier for us to use. They (and we) use a unit for counting the number of atoms in a sample of matter. This unit is called the MOLE. One mole is equal to 6.022×10^{23} atoms of any element. Using the term "mole" is similar to using the word "dozen". We all know that a dozen means 12. Well, a mole means 6.022×10^{23}.

Example 1: 3.67×10^{24} molecules NCl_3 = ? moles

Example 2: 0.006972 moles H_2O = ? molecules

Now, that is all well and good. But we, as humans, still cannot see individual atoms. How can we count them? The chemists came up with a way that we could measure moles. Chemists decided that one mole of any element was equal to the average atomic mass of that element in grams. They call it the "molar mass" of the element. For example, the average atomic mass of carbon on the Periodic

Table is 12.011. Therefore, 12.011 grams of carbon is equal to one mole of carbon. And, one mole of carbon is equal to 6.022×10^{23} atoms of carbon. So, 6.022×10^{23} atoms of carbon have a mass of 12.011 grams.

Example 3: 365.4 g P_2O_5 = ? moles

Example 4: 1.375 moles $Ca(NO_3)_2$ = ? grams

Conversion Factors to use:

1 mole = 6.022×10^{23} atoms or molecules
= molar mass of element or compound
"molar mass" = decimal # on P.T. in grams

This means that when converting among "moles", "grams", "atoms", and "molecules"...
the number "1" ALWAYS goes in front of the unit "moles"
the number "6.022×10^{23}" ALWAYS goes in front of the unit "atoms"
the number "6.022×10^{23}" ALWAYS goes in front of the unit "molecules (mcs)"
the "molar mass" (decimal # on Periodic Table) ALWAYS goes in front of the unit "grams"

Example 5: 3.67×10^{24} mcs PI_3 = ? grams

Example 6: 359 g $Ni(ClO)_2$ = ? mcs

MOLES AND MOLECULES WORKSHEET

Part 1 - Determine the number of moles in each of the quantities below.

1.) 2.50×10^{23} molecules of NaCl

2.) 1.25×10^{23} molecules of H_2SO_4

3.) 1.00×10^{24} molecules of $KMnO_4$

4.) 7.4×10^{24} molecules of KCl

5.) 3.5×10^{25} molecules of $CuSO_4$

Part 2 - Determine the number of molecules in each of the quantities below.

6.) 2.5 moles of NaCl

7.) 0.50 moles of H_2SO_4

8.) 1.70 moles of $KMnO_4$

9.) 0.25 moles of KCl

10.) 3.2 moles of $CuSO_4$

ANSWERS:

1.) 0.415 moles 2.) 0.208 moles 3.) 1.66 moles 4.) 12 moles 5.) 58 moles
6.) 1.5×10^{24} mcs 7.) 3.0×10^{23} mcs 8.) 1.02×10^{24} mcs 9.) 1.5×10^{23} mcs 10.) 1.9×10^{24} mcs

MOLES AND MASS WORKSHEET

Part 1 - Determine the number of moles in each of the quantities below.

1.) 25.0 grams of NaCl

2.) 125 grams of H_2SO_4

3.) 100. grams of $KMnO_4$

4.) 74.5 grams of KCl

5.) 35 grams of $CuSO_4$

Part 2 - Determine the mass (number of grams) in each of the quantities below.

6.) 2.5 moles of NaCl

7.) 0.50 moles of H_2SO_4

8.) 1.70 moles of $KMnO_4$

9.) 0.25 moles of KCl

10.) 3.2 moles of $CuSO_4$

ANSWERS:

1.) 0.427 moles	2.) 1.27 moles	3.) 0.633 moles	4.) 0.999 moles	5.) 0.22 moles
6.) 150 g	7.) 49 g	8.) 269 g	9.) 19 g	10.) 510 g

PRACTICE WITH MOLE CONVERSIONS WORKSHEET

1. How many moles are equal to 2.548 grams of boron trifluoride, BF_3?

2. How many grams are there in 2.45×10^{24} molecules of ammonia, NH_3?

3. How many moles are equal to 5.29×10^{23} atoms of carbon?

4. What is the mass (in grams) of 6.759 moles of sodium chloride, NaCl?

5. How many molecules of H_3PO_4 are contained in 0.257 moles of H_3PO_4?

6. How many atoms of copper are equal to 49.5 grams of copper?

7. What is the mass (in grams) of 7.14×10^{23} molecules of $C_6H_{12}O_6$?

8. How many moles are equal to 8.392×10^{23} atoms of uranium?

9. What is the mass (in grams) of 5.685 moles of sodium bicarbonate, $NaHCO_3$?

10. How many molecules are equal to 0.027 moles of calcium carbonate, $CaCO_3$?

11. How many moles are equal to 93.75 grams of sodium sulfate, Na_2SO_4?

12. How many molecules are equal to 103.74 grams of lead nitrate, $Pb(NO_3)_2$?

*13. A large piece of aluminum foil has a mass of 35.25 grams. What mass of pure tin would contain the same number of atoms as the aluminum foil?

1.) 0.03758 moles 2.) 69.2 grams 3.) 0.878 moles 4.) 395.4 grams
5.) 1.55×10^{23} molecules 6.) 4.69×10^{23} atoms 7.) 213 grams 8.) 1.394 moles
9.) 477.5 grams 10.) 1.6×10^{22} molecules 11.) 0.6597 moles
12.) 1.8862×10^{23} molecules *13.) 155.0 grams

UNIT 3 REVIEW WORKSHEET

Part 1 - Fill in the blanks

Atoms are made up of electrons, which have a __1__ charge; __2__, which have a positive charge; and __3__, which are neutral. The latter two particles are found in the __4__ of the atom. It was __5__ who discovered the nucleus of the atom. The nucleus has a __6__ charge and it occupies a very small volume in the atom. In contrast, the negatively-charged __7__ occupies most of the volume of the atom. The number of __8__ in the nucleus of the atom is the atomic __9__ of that element. Because the atom is electrically neutral, the number of protons and __10__ are equal. The sum of the __11__ and neutrons is the mass number. Atoms of the same element are identical in most respects, but they can differ in the number of __12__ in the nucleus. Atoms that have the same number of protons but different mass numbers are called __13__. The __14__ of an element is the weighted average of the masses of the isotopes of that element. Two isotopes of sulfur are ^{32}S and ^{34}S. An atom of the sulfur-32 isotope contains __15__ protons and __16__ neutrons. The sulfur-34 isotope has __17__ protons and __18__ neutrons. Each of the three known isotopes of hydrogen has __19__ proton(s) in the nucleus. The most common hydrogen isotope has __20__ neutrons. It has a mass number of __21__ and is called hydrogen-1.

1. ________________
2. ________________
3. ________________
4. ________________
5. ________________
6. ________________
7. ________________
8. ________________
9. ________________
10. ________________
11. ________________
12. ________________
13. ________________
14. ________________
15. ________________
16. ________________
17. ________________
18. ________________
19. ________________
20. ________________
21. ________________

Part 2 - Solve the following problems.

22. Calculate the average atomic mass of oxygen given the percent abundance of each of its naturally-occurring isotopes: oxygen-16 is 99.76%, oxygen-17 is 0.037%, and oxygen-18 is 0.204%.
23. What is the mass of 6.00 moles of aluminum?
24. How many moles contain 8.53 x 10^{23} atoms of platinum?
25. How many atoms of carbon are in 3.25 moles of carbon?
26. How many moles of SO_2 are in 160.0 grams of SO_2?
27. What is the mass of 3.58 x 10^{24} atoms of nitrogen?
28. How many molecules does 80.0 grams of K_2O contain?

Part 3 - Fill in the following chart.

Name	# of e^-	Atomic #	Mass #	# of n^o	# of p^+	Symbol
arsenic - 76						
		27	58			
				48	36	
						^{11}B
	25		56			
				8	6	
		7		7		

Answers to # 23 - 28:

23.) 162 grams 24.) 1.42 moles 25.) 1.96×10^{24} atoms 26.) 2.496 moles

27.) 83.2 grams 28.) 5.11×10^{23} molecules